TOPIC 01 – QUANTITATIVE CHEMISTRY

IB Chemistry

T01D01

1.1 The mole concept and Avogadro's constant - 2 hours

- 1.1.1 Apply the mole concept to substances (2)
- 1.1.2 Determine the number of particles and the amount of substance (in moles).(3)

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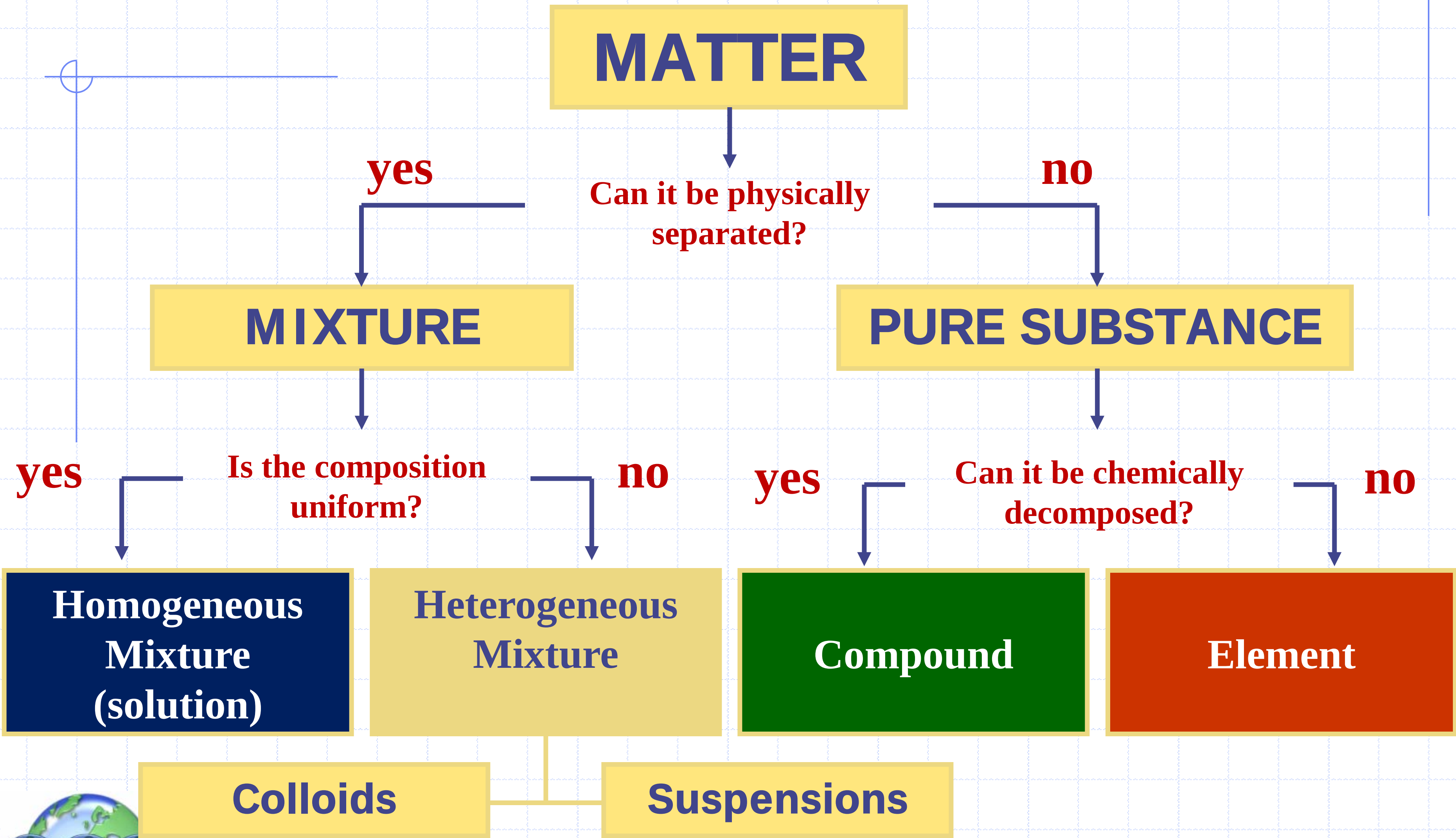
1.1 - Matter Flow Chart



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1.1 - The Mole

- Chemical reactions involve atoms and molecules.
- The ratios with which elements combine depend on the **number of atoms** not on their mass.
- The masses of atoms or molecules depend on the substance.
- Individual atoms and molecules are extremely small. Hence a larger unit is appropriate for measuring quantities of matter.
  - A **mole** is equal to exactly the number of atoms in exactly **12.0000 grams of carbon 12.**
  - This number is known as Avogadro's number.
    - 1 mole is equal to **6.02×10^{23}** particles.

1.1 – Relative Atomic Mass

- The weighted average (A_r) is only given a value compared to the mass of the carbon-12 atom
- 1/12th of carbon 12 would be one unit!
- So it's simply the weighted average divided by one!
- $A_r = \dfrac{weighted\ average\ mass\ of\ isotopes\ of\ the\ element}{\frac{1}{12} x\ the\ mass\ of\ one\ atom\ of\ carbon{-}12}$

1.1 - Definitions of the Mole

- 1 mole of a substance has a mass equal to the **formula mass in grams**.
- *Examples*
  - 1 mole H_2O is the number of molecules in 18.01 g H_2O
  - 1 mole H_2 is the number of molecules in 2.02 g H_2.
  - 1 mole of atoms has a mass equal to the atomic weight in grams.
  - 1 mole of particles = 6.02×10^{23} particles for any substance!
- The **Molar mass** is the mass of one mole of a substance
  - **Avogadro's number** is the number of particles (molecules) in one mole for any substance

1.1 – Calculating Molar Mass

- To find the **molar mass** you can simply add together the **relative atomic mass** of each of the atoms present in a molecule
- You may always round off to two unit passed the decimal for the molar mass
  - Oxygen = 16.00 g mol^{-1}
  - Hydrogen = 1.01 g mol^{-1}
  - Carbon = 12.01 g mol^{-1}
  - etc

1.1 – Calculating Molar Mass

- Calculate the Molar Mass for glucose ($C_6H_{12}O_6$)
  - C – 12.01 x 6
  - H – 1.01 x 12
  - O – 16 x 6
  - (12.01x6) + (1.01x12) + (16x6) = 180.18 g mol^{-1}
- **This value means that for every mole of glucose, it has a mass of 180.18 grams.** OR that there are 180.18 grams of glucose in one mole.

1.1 - Conversions

- This now means that we can make simple conversions.
- We need to keep in mind:
  - UNITS!!!
  - We are only multiplying by a value of 1 when converting
  - UNITS!!!
  - We want certain units to cancel out so we can flip the numerator/denominator as fits the problem

1.1 - Simple Conversions:

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- Mole / Mass Conversions -

Use the Molar Mass of a substance to convert from Moles to Mass and Mass to Moles

Mass to Moles →

80. g $CuSO_4$	1 mol $CuSO_4$
	159.5 g $CuSO_4$

= 0.50 mol $CuSO_4$

$$80.\ g\ CuSO4\ x\ \frac{1\ mol\ CuSO4}{159.5\ g\ CuSO4} = 0.50\ mol\ CuSO4$$

Moles to Mass →

0.50 mol $CuSO_4$	159.5 g $CuSO_4$
	1 mol $CuSO_4$

= 80. g $CuSO_4$

$$0.50\ mol\ CuSO4\ x\ \frac{159.5\ g\ CuSO4}{1\ mol\ CuSO4} = 80.\ CuSO4$$

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1.1 - Simple Conversions:

- Mole / Molecule Conversions -

Use Avogadro's Number : 6.022×10^{23} molecules (mc) in one mole of the substance

Moles to (mc) →

$$\frac{2 \text{ mol } CuSO_4 \;|\; 6.022\text{x}10^{23} \text{ (mc) } CuSO_4}{\;|\; 1 \text{ mol } CuSO_4} = 1.2\text{x}10^{24} \text{ (mc) } CuSO_4$$

(mc) to Moles →

$$\frac{1.2\text{x}10^{24} \text{ (mc) } CuSO_4 \;|\; 1 \text{ mol } CuSO_4}{\;|\; 6.022\text{x}10^{23} \text{ (mc) } CuSO_4} = 2 \text{ mol } CuSO_4$$

1.1 - Simple Conversions:

- Mole / Volume Conversions -

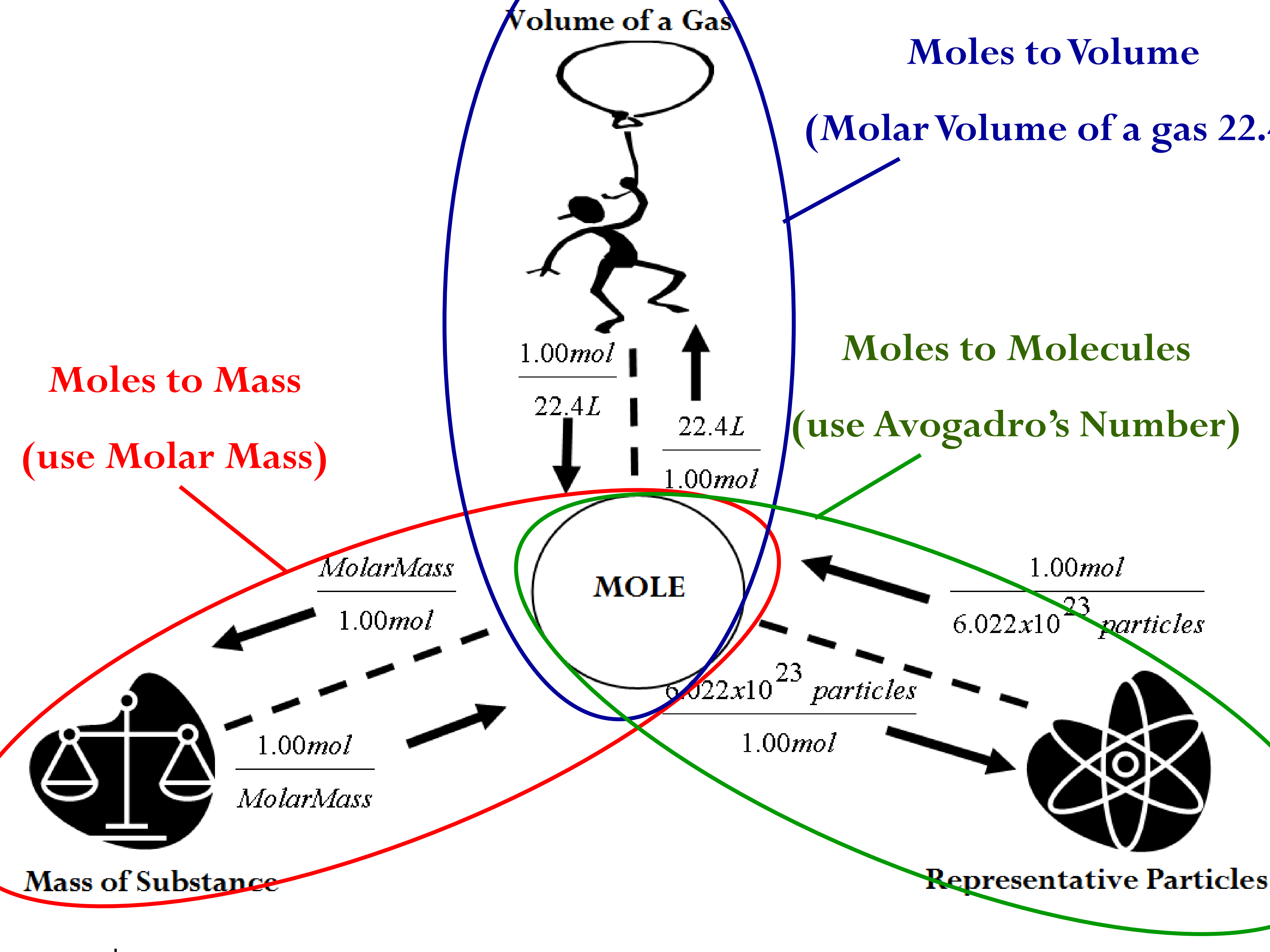
At STP (standard temperature and pressure) one mole of any gas takes up 22.4 L of space

Moles to Volume →

$$\frac{2 \text{ mol } O_2}{} \Bigg| \frac{22.4 \text{ L } O_2}{1 \text{ mol } O_2} = 44.8 \text{ L } O_2$$

Volume to Moles →

$$\frac{44.8 \text{ L } O_2}{} \Bigg| \frac{1 \text{ mol } O_2}{22.4 \text{ L } O_2} = 2 \text{ mol } O_2$$



Volume of a Gas
Moles to Volume
(Molar Volume of a gas 22.
Moles to Molecules
(use Avogadro's Number)
Moles to Mass
(use Molar Mass)
1.00mol / 22.4L
22.4L / 1.00mol
MOLE
MolarMass / 1.00mol
1.00mol / MolarMass
1.00mol / 6.022x10^23 particles
6.022x10^23 particles / 1.00mol
Mass of Substance
Representative Particles

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1.2 Formulas - 3 hours

- 1.2.1 Define the terms relative atomic mass (Ar) and relative molecular mass (Mr). (1) – taught with topic 02
- 1.2.2 Calculate the mass of one mole of a species from its formula. (2)
- 1.2.3 Solve problems involving the relationship between the amount of substance in moles, mass and molar mass. (3)
- 1.2.4 Distinguish between the terms empirical formula and molecular formula. (2) - taught with topic 02
- 1.2.5 Determine the empirical formula from the percentage composition or from other experimental data. (3) – topic 02
- 1.2.6 Determine the molecular formula when given both the empirical formula and experimental data. (3) – topic 02

**We will study formulas with nomenclature soon!!

1.2 - Stoich Problem #1

- If 80.0 grams of copper (II) sulfate reacts with an excess of iron, how many grams of copper will be produced?

Given: (g CuSO4)   Find:(g Cu)

Balanced Eqn: $CuSO_4 + Fe \rightarrow FeSO_4 + Cu$

How to get there: g $CuSO_4$, mol $CuSO_4$, mol Cu , g Cu

80.0 g $CuSO_4$	1 mol $CuSO_4$	1 mol Cu	63.5 g Cu	= 31.8 g Cu
	159.5 g $CuSO_4$	1 mol $CuSO_4$	1 mol Cu	

1.2 - Stoich Problem #2

- The base calcium hydroxide neutralizes 60.0 g nitric acid, how many g of salt will be produced?

Given: (g HNO_3) Find: (g salt)

Balanced Eqn: $2\ HNO_3 + Ca(OH)_2 \rightarrow 2\ H_2O + Ca(NO_3)_2$

How to get there: g HNO_3 , mol HNO_3 , mol $Ca(NO_3)_2$, g $Ca(NO_3)_2$

60.0 g HNO_3	1 mol HNO_3	1 mol $Ca(NO_3)_2$	160 g $Ca(NO_3)_2$	=	76.2 g $Ca(NO_3)_2$
	63 g HNO_3	2 mol HNO_3	1 mol $Ca(NO_3)_2$		

1.2 - Stoich Problem #3

- 2.0×10^{25} atoms of copper is reacted with an excess of a silver (I) nitrate solution, how many grams of silver are produced?

Given: (atoms Cu) Find:(g Ag)

Balanced Eqn: $Cu + 2\ AgNO_3 \rightarrow 2\ Ag + Cu(NO_3)_2$

How to get there: atoms Cu , mol Cu , mol Ag , g Ag

$$\frac{2.0\times10^{25}\ (a)\ Cu}{} \left| \frac{1\ mol\ Cu}{6.022\times10^{23}\ (a)\ Cu} \right| \frac{2\ mol\ Ag}{1\ mol\ Cu} \left| \frac{107.9\ g\ Ag}{1\ mol\ Ag} \right. = 7.2\times10^{3}\ g\ Ag$$

1.2 - Stoich Problem #4

- An excess of methane gas combusts in the presence of 1.00 kL of oxygen, what volume of carbon dioxide will be produced at STP?

Given: (1 kL O_2) Find: (L CO_2)

Balanced Eqn: $CH_4 + 2\ O_2 \rightarrow CO_2 + 2\ H_2O$

How to get there: kL O_2 , L O_2 , mol O_2 , mol CO_2 , L CO_2

1.00 kL O_2	1000 L O_2	1 mol O_2	1 mol CO_2	22.4 L CO_2	= 500 L CO_2
	1 kL O_2	22.4 L O_2	2 mol O_2	1 mol CO_2	

1.3 Chemical equations – 1 hour

- 1.3.1 Deduce chemical equations when all reactants and products are given. (3)
- 1.3.2 Identify the mole ratio of any two species in a chemical equation. (2) – covered in 1.2
- 1.3.3 Apply the state symbols (s), (l), (g) and (aq). (2)

1.3 – Chemical Equations

- Can be expressed in three methods
  - Work equations
    - Iron + chorine → iron (III) chloride
  - Skeletal equations
    - $Fe + Cl_2 \rightarrow FeCl_3$
  - Balanced equations
    - $2Fe + 3Cl_2 \rightarrow 2FeCl_3$
    - (by adding coefficients, not subscripts)
    - The **law of conservation of mass** states that matter cannot be created nor destroyed, thus there must be an equal number of atoms in the **reactants** and **products**. (this is not true of moles)

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1.3 – Chemical Equations

- Other important information may also be given:
  - Mole Ratio:
    - Ratio of moles of one compound to another, the coefficients provide this
  - State symbols
    - $2Fe(s) + 3Cl_2(g) \rightarrow 2FeCl_3(s)$
  - Thermodynamic data
    - $2Fe + 3Cl_2 \rightarrow 2FeCl_3 \quad \Delta H^o = -750\ kJ\ mol^{-1}$
- Keep in simplest form:
  - $4H_2(g) + 2O_2(g) \rightarrow 4H_2O(l)$ vs...
  - $2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$